**Oxidation-Reduction (Redox)**

An **oxidation-reduction** or **redox reaction** is a reaction that involves the transfer of electrons between chemical species (the atoms, ions, or molecules involved in the reaction). Some examples of common redox reactions are shown below.

CH4(g) + 2 O2(g) → CO2(g) + 2 H2O(g) (combustion of methane)

2 Cu(s) + O2(g) → 2 CuO(s) (oxidation of copper)

During a redox reaction, some species undergo **oxidation**, or the loss of electrons, while others undergo **reduction**, or the gain of electrons.

For example, consider the reaction between iron and oxygen to form rust:

4 Fe(s) + 3 O2(g) → 2 Fe2O3(s) (rusting of iron)

In this reaction, **Fe** loses electrons to form **Fe3+ ions**

and **O2** gains electrons to form **O2−** **ions.**

In other words, **iron is *oxidized,*** and **oxygen is *reduced***.

To help you memorize this, remember the pneumonic device, ***OIL RIG***

**O**xidation **R**eduction

**I**s **I**s

**L**osing electrons **G**aining electrons

Note that when a species gets oxidized, its oxidation number goes up and

when a species gets reduced, its oxidation number goes down.

But what is an oxidation number?

**Rules for Oxidation numbers**

An atom’s **oxidation number** (or **oxidation state**) is the imaginary charge that the atom would have if all of the bonds to the atom were completely ionic. Oxidation numbers can be assigned to the atoms in a reaction using the following guidelines:

1. **An atom of a free element** has an oxidation number of 0. For example, each atom in O2 has an oxidation number of 0. The same is true for each in H2​, each S atom in S8, each C atom in C and so on.
2. **A monatomic ion** has an oxidation number equal to its charge. For example, the oxidation number of Cu2+ is +2. And the oxidation number of Br−  is -1.
3. When combined with other elements, **alkali metals** (Group 1) always have an oxidation number of +1, while **alkaline earth metals** (Group 2) always have an oxidation number of +2.
4. **Fluorine** has an oxidation number of -1in all compounds.
5. **Hydrogen** has an oxidation number of +1 in most compounds. The major exception is when hydrogen is combined with metals, as in NaH, or CaH2. In these cases, the oxidation number of hydrogen is -1.
6. **Oxygen** has an oxidation number of -2 in most compounds. The major exception is in peroxides (compounds containing O22−, where oxygen has an oxidation number of -1. Examples of common peroxides include H2O2 and Na2O2.
7. The **other halogens** (Cl, Br and I) have an oxidation number of -1, in compounds, unless combined with oxygen or fluorine. For example, the oxidation number of Cl in the ion ClO4− is +7 (since O has an oxidation number of -2, and the overall charge on the polyatomic ion is -1. *Explanation:* *Each O is -2. 4 x -2 = -8. Therefore, Cl needs to be +7 because +7 -8 = -1 (-1 is the charge on ClO4−)*
8. **The sum of the oxidation numbers** for all atoms in a neutral compound is equal to zero, while the sum for all atoms in a polyatomic ion is equal to the charge on the ion.

**Example 1: Assigning oxidation numbers**

**What is the oxidation number of each atom in:**

**(a) SF6**

**(b) H3PO4**

**(c) IO3−**

**S+6F-16  (+6 + 6(-1) = 0)**

**H+13P+5O-24 (3(+1) + (+5) + 4(-2) = 0)**

**I+5O-23−  ((+5) + 3(-2) = -1)**

**Recognizing redox reactions**

How do we actually use oxidation numbers to identify redox reactions? To find out, let’s revisit the reaction between iron and oxygen, this time assigning oxidation numbers to each atom in the equation:

4 Fe**0**(s) + 3 O2**0**(g) → 2 Fe**+3**2O**-2**3(s) (rusting of iron)

Iron changes from an oxidation number of 0 to an oxidation number of +3. Oxygen changes from an oxidation number of 0 to an oxidation number of -2.

We can conclude that oxidation involves an *increase* in oxidation number, while reduction involves a *decrease* in oxidation number.

**Oxidizing Agents and Reducing Agents**

An **oxidizing agent** is a substance that causes oxidation by accepting electrons; therefore, it gets **reduced.** A **reducing agent** is a substance that causes reduction by losing electrons; therefore, it gets **oxidized**.

**Writing and Balancing Redox Reactions**

In the **ion**-**electron method** (also called the **half**-**reaction method**), the **redox equation** is separated into two **half**-**equations** - one for oxidation and one for reduction. Each of these **half**-**reactions** is balanced separately and then combined to give the balanced **redox equation**.

**HOW TO WRITE AND BALANCE A REDOX EQUATION**

We will learn to balance a redox reaction using the following reaction as an example:

**HCl + KMnO4 🡪 H2O + KCl + MnCl2 + Cl2**

**Step 1)**  Find the **Oxidation Numbers** for each element.

H**+1**Cl**-1** + K**+1**Mn**+7**O**-2**4 ------> H**+1**2O**-2** + K**+1**Cl**-1** + Mn**+2**Cl**-1**2 + Cl**0**2

**Step 2)**  Find out which element is being oxidized and which element is being reduced; then

start to write out the 1/2 rxns for each.

Oxidation: Cl**-1** 🡪 Cl**0**2

Reduction: Mn**+7** 🡪 Mn**+2**

**Step 3)** Then Balance both the Number of atoms and the Charge...

Oxidation: 2 Cl**-1** 🡪 Cl**0**2 + 2 e**-1**

Reduction: 5 e**-1** + Mn**+7** 🡪 Mn**+2**

**Step 4)** Balance out the electrons so that the number of electrons gained = the number lost.

Oxidation: (2 Cl**-1** 🡪 Cl**0**2 + 2 e**-1**) x 5 = 10 Cl**-1** ------> 5 Cl**0**2 + 10 e**-1**

Reduction: (5 e**-1** + Mn**+7** 🡪 Mn**+2**) x 2 = 10e**-1** + 2 Mn**+7** ------> 2 Mn**+2**

**Step 5)** Now add the two reactions together like a math problem, canceling out the electrons.

Oxidation: 10 Cl**-1** 🡪 5 Cl**0**2 + ~~10 e~~**-1**

Reduction: ~~10e~~**-1** + 2 Mn**+7** 🡪 2 Mn**+2**

**-----------------------------------------------------------**

**10** Cl**-1** + **2** Mn**+7** 🡪 **2** Mn**+2** + **5** Cl**0**2

**Step 6)** Reinsert the coefficients into the original equation.

H**+1**Cl**-1** + K**+1**Mn**+7**O**-2**4 🡪 H**+1**2O**-2** + K**+1**Cl**-1** + Mn**+2**Cl**-1**2 + Cl**0**2

becomes.....

**10**  HCl + **2** KMnO4 🡪 H2O + KCl + **2** MnCl2 + **5** Cl2

**Step 7)** Now check to make sure everything is balanced, and if not, the try to rebalance it.

Since the K’s were not balanced, The KClis multiplied by **2**.

**10**  HCl + **2** KMnO4 🡪 H2O + **2**  KCl + **2** MnCl2 + **5** Cl2

Now to balance the Cl’s (since there are 16 on the right, we need 16 on the left.)

**16** HCl + **2** KMnO4 🡪 H2O + **2** KCl + **2** MnCl2 + **5** Cl2

Finally, to balnce the H’s (and also the O’s) 8 H2O’s are needed.

**16** HCl + **2** KMnO4 🡪 **8** H2O + **2** KCl + **2** MnCl2 + **5** Cl2

**Teacher:** Mr. Baruch **Unit:** Redox

**Subject:** Chemistry

Name: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ **Date:** \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**Redox**

**1.** What is oxidation? Give an example of this process.

**2.** What is reduction? Give an example of this process.

**3.** What is meant by a “redox reaction”?

**4.** In the following reactions, identify which elements are being oxidized, and which are being reduced. Show the oxidation numbers of the oxidized and reduced elements:

**a.** 2 K (s) + H2SO4 (aq) ----> K2SO4 (aq) + H2 (g)

**b.** H2 (g) + CuO (s) -----> Cu (s) + H2O (l)

**5.** For the equations above, write the electronic equations for the elements oxidized and reduced in each.

**Teacher:** Mr. Baruch **Unit:** Redox

**Subject:** Chemistry

Name: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ **Date:** \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

# Balancing Half Reactions Worksheet

**Use the half-reaction method to balance each of the following equations. Show each half reaction.**

**1.** Li + H2O ----> LiOH + H2

**2.** Al + HCl ------> AlCl3 + H2

**3**. H2S + HNO3 -----> H2SO4 + NO2 + H2O

**4.** Cu + HNO3 -----> Cu(NO3)2 + NO + H2O

**5.** KClO3 ------> KCl + O2

ELECTROCHEMICAL CELL

A device that produces useable electric energy

from a **spontaneous** chemical rxn;  *a battery*.



Zn**0(s)**  Zn2+ + 2e- Cu2+ + 2e- Cu**0(s)**

oxidation reduction

Anode Cathode

- is negative - is positive

- is where oxidation occurs - is where reduction occurs

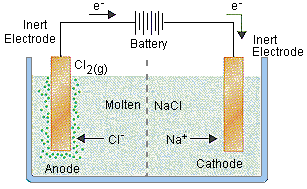
- e- flow from anode - e- flow to the cathode

- e- generated by the - e- accumulated and used

oxidation 1/2 rxn to make reduction happen

**ELECTROLYTIC CELL**

A device that drives a **non-spontaneous** chemical rxn by using an external electrical source. (It uses a battery)



2 Cl- Cl2(g) + 2e- 2 Na+ +2e- 2 Na0(s)

Anode Cathode

- is positive - is negative

- oxidation occurs at anode - reduction occurs at cathode

- neg. ions attracted to anode; - positive ions attracted to

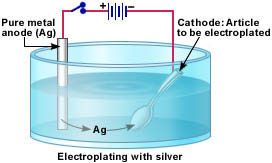
- electrons produced here. cathode, battery supplies

electrons to cathode.

ELECTROPLATING

Electric current used to deposit a layer of of metal, such as silver, on the object to be plated.

Same system as the electrolytic cell.



Ag Ag+ + 1e- Ag+ + 1e- Ag

Anode Cathode

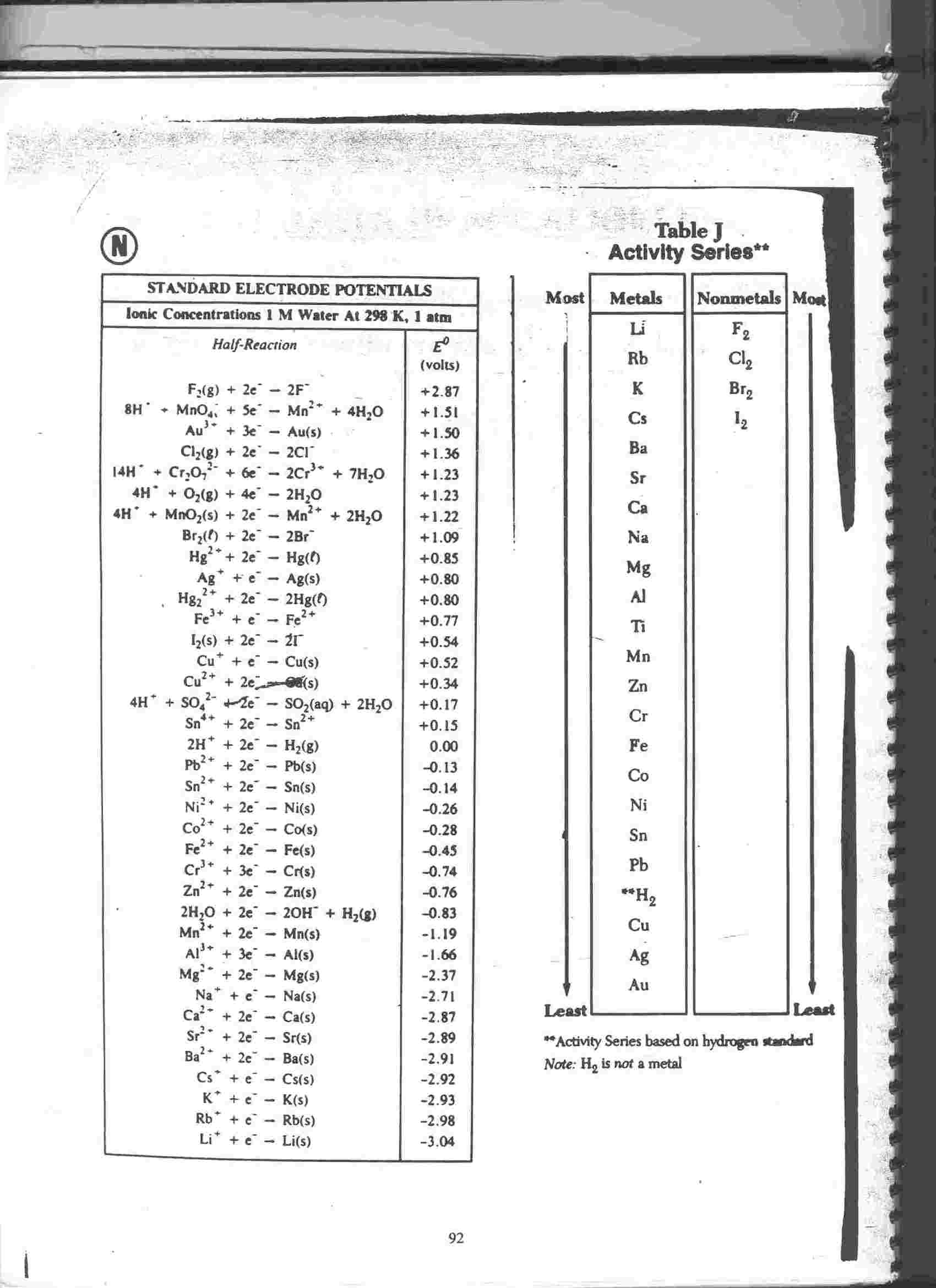
- is positive - is negative

- oxidation occurs at anode - reduction occurs at cathode

- positive ions provided - positive ions attracted to

by anode cathode

- electroplating occurs at cathode



Name: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ **Date:** \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**Electrochemistry**

**1.** Write the word **anode**  or **cathode** next to each line.

**a.** The electrode at which oxidation occurs. \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**b.**The electrode at which reduction occurs. \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**c.**The negative electrode in an electrochemical cell. \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**d.**The negative electrode in an electrolytic cell. \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**e.**In an electrolytic cell, the electrode that attracts Cl- ions . \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**f.**The part of an electroplating system that is provided by a fork. \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**g.**The part of an electroplating system that generates Ag+ ions. \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**2.** Define the term reducing agent.

**3.** Define the term oxidizing agent.

**4.** The most active reducing agent among the elements is:

**a.** iodine **b.** cesium **c.** fluorine **d.** lithium

Explain why: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**5.** The most active oxidizing agent among the elements is:

**a.** iodine **b.** cesium **c.** fluorine **d.** lithium

Explain why: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**6.** How do anodes and cathodes differ in electrochemical cells vs. electrolytic cells?

**OVER 🡪**

**7.** Write out the **net ionic (or skeletal redox) equation** for an electrochemical cell that has

MgCl2 in one 1/2 cell, and SnCl2 in the other 1/2 cell.

**8.** Draw the full electrochemical cell diagram for question #7 on the back of this sheet.

**Class:**  Chemistry

Name: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ **Date:** \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

# Electrochemical Cell Worksheet

You are given two beakers. In Beaker #1 is a solution of lead nitrate ( Pb(NO3)2 ) and a piece of solid lead metal. In Beaker #2 is a solution of nickel nitrate ( Ni(NO3)2 ) and a piece of solid nickel metal. The two pieces of metal are connected by wires to each other through a voltmeter which reads electrode potential (Ecell) in volts. Finally, a salt bridge is placed between the two beakers. The salt bridge contains a solution of potassium nitrate.

1. Draw a diagram of the picture mentioned above.

1. Which beaker ( half-cell) is undergoing oxidation and which is undergoing

reduction?

3. Write the 1/2 reactions for oxidation and reduction of these two half-cells.

4. What does the **Ecell**  equal for this reaction?

5. Which electrode is the anode and which is the cathode?

6. Which electrode is positive and which is negative?

7. In which direction are the electrons flowing (from what to what)?

1. Why will this reaction occur spontaneously?
2. What is the function of the salt bridge?
3. What will happen to the Nickel metal over time, and what will happen to the Lead metal over time?
4. Which of the two solutions will have more metal ions in it over time and which will have less over time?